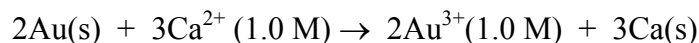


- Calculate the standard free-energy change for the following reaction at 298 K.



Marks
2

The reduction half cell reactions and E^0 values are:



In the reaction, Au is being oxidized and so the overall cell potential is:

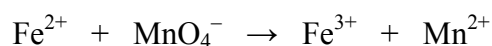
$$E^0 = ((-2.87) - (+1.50)) \text{ V} = -4.37 \text{ V}$$

The reaction involves 6 electrons so, using $\Delta G^0 = -nFE^0$:

$$\begin{aligned} \Delta G^0 &= - (6) \times (96485 \text{ C mol}^{-1}) \times (-4.37 \text{ V}) = +2530000 \text{ J mol}^{-1} \\ &= +2.53 \times 10^3 \text{ kJ mol}^{-1} \end{aligned}$$

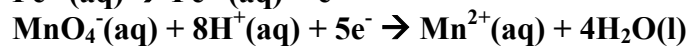
Answer: $+2.53 \times 10^3 \text{ kJ mol}^{-1}$

Complete and balance the following equation for the reaction between iron(II) ions and permanganate ions in an acidic solution.

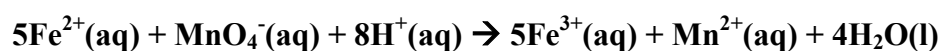


2

In acid, the relevant half cells are:

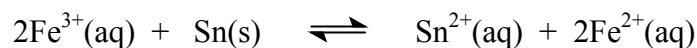


Giving an overall reaction:

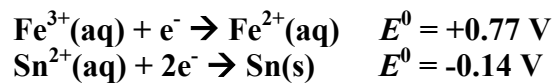


ANSWER CONTINUES ON THE NEXT PAGE

- What is the value of the equilibrium constant for the following reaction at 298 K?



The reduction half cell reactions and E^0 values are:



In the reaction, Sn is being oxidized and so the overall cell potential is:

$$E^0 = ((+0.77) - (-0.14)) \text{ V} = +0.91 \text{ V}$$

The reaction involves 2 electrons so, using $E^0 = \frac{RT}{nF} \ln K$:

$$\ln K = E^0 \times \frac{nF}{RT} = (+0.91 \text{ V}) \times \left(\frac{2 \times 96485 \text{ C mol}^{-1}}{8.314 \text{ J K}^{-1} \text{ mol}^{-1} \times 298 \text{ K}} \right) = 70.9$$

$$K = e^{70.9} = 6.05 \times 10^{30}$$

Answer: 6.05×10^{30}